the carbon–fluorine bond dissociation energy in vinyl fluoride, representing a rise of 19 kcal. mole^{-1} in carbon–fluorine bond energy in changing from sp^3 hybridized carbon in methyl fluoride to sp^2 hybridized carbon in vinyl fluoride. This also serves to illustrate that the rearrangement process leading to formation of $C_2H_3F^+$ proceeds with only a small, if not zero, activation energy.

Ions produced by loss of fluorine from the parent ion give peaks which are intense in all three spectra. These ions can be formed either by loss of a neutral fluorine atom or by a pair production process involving formation of a negative F^- ion. The relatively high appearance potentials for these ions indicate the former process as the more likely and, in fact, the appearance potential curve for $C_2H_5CF_2^+$ from 1,1,1-trifluoropropane shows an indication of divergence to a lower value at the foot of the curve. Unfortunately, the very low intensity at this point makes it impossible to locate the lower limit. In only one case was it possible to measure a checking value: the $C_2H_3CF_2^+$ ion has been measured for both 3,3,3-trifluoropropene and 1,1,1-trifluoropropane. The appearance potential of the ion from the two sources can be correlated by assuming these two processes for formation

$$C_2H_3CF_3 + e \longrightarrow C_2H_3CF_2^+ + F + 2e \tag{12}$$

and

$$C_2H_5CF_3 + e \longrightarrow C_2H_3CF_2^+ + HF + H + 2e \tag{13}$$

Assuming (12), the calculated appearance potential for (13) is 13.2$_5$ e.v., compared to the observed 13.6 e.v. No other combination of the many possible processes predicts the appearance potential to better than 1.0 e.v. These results indicate that negative ion formation does not play an important role in the decomposition by electron impact of the three fluorocarbons examined in this study.[26]

It is evident from the foregoing discussion that the values chosen for the heats of formation of the trifluoromethyl radical and its positive ion lead to heats of formation of the R–CF_3 compounds which satisfactorily interpret the appearance potentials of many of the positive ions. Processes 5, 6, 7 and 8 are cases where the correct process is unambiguous and the appearance potentials can be calculated with some accuracy. In all four of these cases the calculated value is somewhat above that observed but only by a few kcal. mole^{-1}. A slightly better fit might be obtained by taking $\Delta H_f(CF_3^+) = 127$ kcal. mole^{-1} and $\Delta H_f(CF_3\cdot) = -107$ kcal. mole^{-1}, but the present results alone do not warrant this change.

[CONTRIBUTION FROM THE DEPARTMENT OF BIOLOGICAL CHEMISTRY, HARVARD MEDICAL SCHOOL, AND THE CANCER RESEARCH INSTITUTE OF THE NEW ENGLAND DEACONESS HOSPITAL, BOSTON, MASSACHUSETTS]

The Elbs Peroxydisulfate Oxidation: Kinetics[1]

BY E. J. BEHRMAN AND PAMELA P. WALKER

RECEIVED MARCH 15, 1962

The reaction of persulfate ion with 2-hydroxypyridine and *o*-nitrophenol has been studied. The kinetics are first order in both persulfate and phenolate ion. There is a general and positive salt effect and no inhibition by the sulfate radical trapping agent, allyl acetate. A bimolecular reaction of persulfate and the phenolate ions is postulated, while heterolysis or homolysis of persulfate prior to attack by the phenolate ion is considered unlikely.

K. Elbs reported in 1893 that the reaction of *o*-nitrophenol in alkaline solution with persulfate[2] yielded 2,5-dihydroxynitrobenzene (nitroquinol) after subsequent heating in acid solution.[3] Later it was found[4] that the initial product in the reaction was a *p*-hydroxyaryl sulfate which could be cleaved by acid hydrolysis to yield the dihydroxy compound. The reaction has since been extended to include a large number of phenols. The position *para* to the existing phenolic group is substituted preferentially but not exclusively; *ortho*-substitution has been detected in a number of cases.[5,6] *Ortho*-substitution is exclusive when the *para*-position is blocked; no case of *meta*-substitution has been reported. Sethna[7] has reviewed the literature through 1950 and, since that time, a few papers have appeared.[6,8] The extension of this reaction to aromatic amines[9] merits particular mention. It should be noted that in this latter case the substitution is preferentially *ortho*.

The mechanism of the reaction has not been investigated. Baker and Brown[10] postulated an attack on the phenolate ion by the sulfate radical. This hypothesis is supported only indirectly by the fact that persulfate is known to be able to decompose *via* a free radical route[11] and by the fact that coupling products have been found as products of the reaction of persulfate with phenols and aromatic amines.[9e,12]

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[1] This investigation was supported by United States Public Health Service Grant A567 and by United States Atomic Energy Commission Contract AT(30-1)-901 with the New England Deaconess Hospital.

[2] Persulfate = peroxydisulfate = peroxodisulfate. See footnote 4 of ref. 22.

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[4] German Patent 81,068. From P. Friedlaender, *Fortschritte der Theerfarbenfabrikation*, 4th part, 1894–7, p. 126.

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Experimental

Materials: 2-Hydroxypyridine.—The material obtained from the Aldrich Chemical Co. was distilled *in vacuo* (b.p. 181 to 185° (24 mm.)) and then recrystallized from carbon tetrachloride or benzene to yield white crystals m.p. 105 to 107°, $\epsilon_{293\,\mathrm{m}\mu}^{\mathrm{water}}$ 5900. The distillation was found necessary to remove a persistent colored impurity which could not be removed by crystallization alone.

o-Nitrophenol.—The Eastman or Matheson, Coleman Bell product was recrystallized from 95 per cent. ethanol. Allyl acetate was purchased from Eastman and redistilled before use as recommended by Kolthoff, *et al.*[13] Potassium persulfate (J. T. Baker Analyzed) was used without recrystallization since preliminary experiments showed no effect on the rate of the reaction by added Cu^{++} or Fe^{++} ions.

Methods: (a) Persulfate Determination.—Reactions were carried out in a thermostated water bath within ± 0.1° of the indicated temperature. Reactions were generally followed by measuring persulfate disappearance by a modification of the iodometric method of Kolthoff and Carr.[14] Suitable aliquots of the alkaline reaction mixture were pipetted into a slight excess of 0.5 *N* hydrochloric acid to which 4 g. of potassium iodide per 30 ml. of solution had been added just previously. (Persulfate and undissociated phenol do not react at an appreciable rate. See Results.) The flask was stoppered, allowed to stand 10 minutes in the dark and then titrated with standardized thiosulfate. The blank correction was in the range of 0.2 to 0.4 ml. 0.01 *M* thiosulfate.

(b) Measurement of Product Formation.—Neither 2-hydroxypyridine nor *o*-nitrophenol, in contrast to most phenols, reduces the Folin phenol reagent. By reaction with alkaline persulfate, these compounds form, respectively 2-hydroxypyridine-5-sulfate and 2-nitrophenol-4-sulfate which yield, following acid hydrolysis, 2,5-dihydroxypyridine[6] and nitroquinol (2,5-dihydroxynitrobenzene).[3] Some *ortho*-substitution also takes place, but the *ortho*-dihydroxy compounds give the same color yield as the *para*-dihydroxy compounds.[6] These diphenols both reduce the Folin reagent. Aliquots of the reaction mixture were pipetted into equal volumes of 6 *N* hydrochloric acid. Hydrolysis was carried out by heating at 100° for 30 to 40 min. 2,5-Dihydroxypyridine or nitroquinol was determined by the procedure previously described.[15] The optical density of 0.1 μmole of either phenol in the standard assay was 0.160 (2,5-dihydroxypyridine) and 0.175 (nitroquinol).

Results

1. Stoichiometry.—Table I shows the effect of varying the ratio of persulfate to 2-hydroxypyridine on the yield of the product. It is clear that a large excess of phenol must be present to ensure a good yield of the desired product. Two experiments showing a comparable yield of nitroquinol from *o*-nitrophenol are also given.

TABLE I

YIELD OF DIPHENOL

Temp., °C.	(Phenol) *M*	(OH^-) *M*	$(S_2O_8^=)$ *M*	(Phenol)/$(S_2O_8^=)$	Yield, %[a]
29.8	0.2[b]	1.7	0.02	10	74.5[c] ±2
29.8	.5[b]	1.0	.025	20	85[d] ± 3
29.8	.4[b]	2.0	.02	20	85[e]
29.8	.1[b]	0.2	.005	20	89[e]
29.8	1.0[b]	1.0	.025	40	82[c] ± 1
29.8	0.15[b]	1.0	.15	1	56[e]
55	.15[b]	1.0	.15	1	54[d] ± 4
29.8	.03[b]	1.0	.15	0.2	14[c] ± 1
29.8	.2[f]	[g]	.02	10	78[c] ± 1

[a] Based on the smaller component. [b] 2-Hydroxypyridine. [c] Average of two experiments. [d] Average of three experiments. [e] Single determination. [f] Nitrophenol. [g] 0.6 *M* carbonate buffer, pH 9.3.

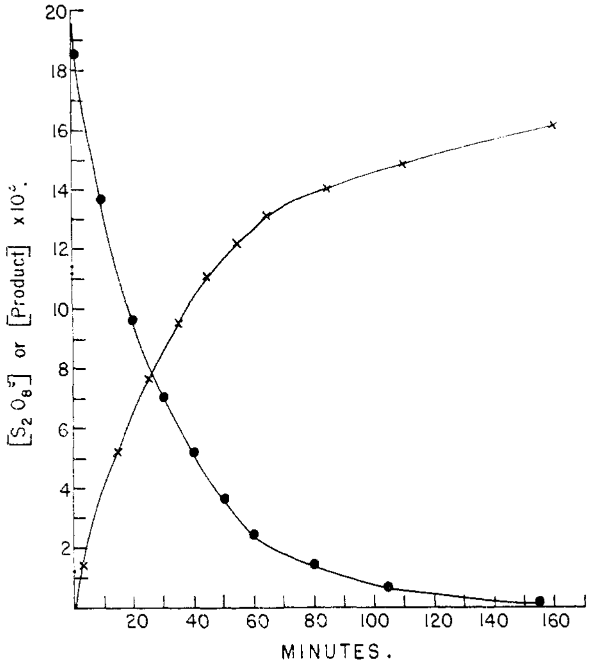

Fig. 1.—The concentration of persulfate and of 2-hydroxypyridine-5-sulfate (as 2,5-dihydroxypyridine) as a function of time: X—X (2-hydroxypyridine-5-sulfate); ●—● $(S_2O_8^=)$, conditions: 0.4 *M* 2-hydroxypyridine, 0.0194 *M* $S_2O_8^=$, 2 *M* NaOH, 29.9°.

Figure 1 shows the results of an experiment in which persulfate disappearance and product appearance were followed simultaneously. A plot of persulfate consumed *vs.* product formed is linear for the major portion of the reaction. The data of Fig. 1 may be expressed in the form: 1.2 (Prod) = $(S_2O_8^={}_{\mathrm{consumed}})$ + 0.0013.

2. Peroxydisulfate Dependence.—When the data of Fig. 1 are plotted as the logarithm of persulfate concentration *vs.* time, a straight line is obtained (in this experiment over at least four half-times). Thus, the consumption of persulfate follows the first order rate expression: $-\,d(S_2O_8^=)/dt = k'(S_2O_8^=)$.

3. Phenol Dependence.—Since reasonable stoichiometry was obtainable only at persulfate–phenol ratios less than 1:10, it would have been equivocal to determine the order in phenol by either (a) measuring the over-all order using equimolar concentrations of reactants or (b) by measuring the dependence on phenol in the presence of excess persulfate. The dependence on phenol concentration was determined by measuring the pseudo first-order constants for persulfate disappearance in the presence of a tenfold excess of phenol over a tenfold initial concentration range. The results of this series of experiments are given in Table II.

From the data of Table II and Fig. 1, it is apparent that the reaction is first order in both phenol and persulfate

$$\frac{-d(S_2O_8^=)}{dt} = k(S_2O_8^=)(ArO^-) \text{ or}$$

$$\frac{-d(S_2O_8^=)}{dt} = k'(S_2O_8^=) \text{ where } k' = k(ArO^-)$$

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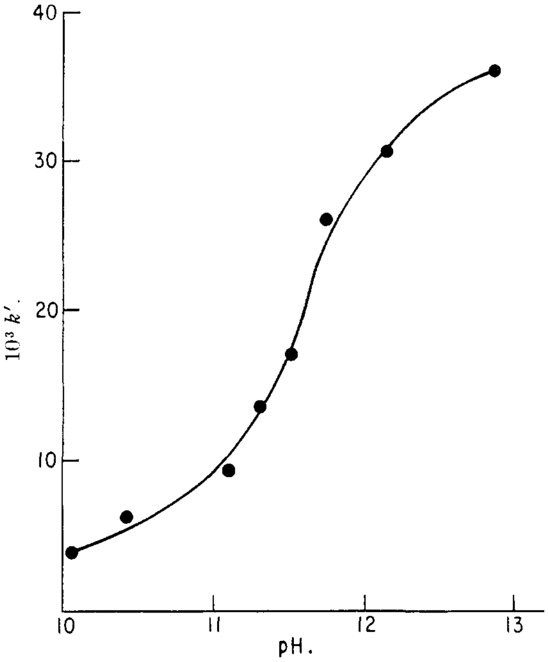

Fig. 2.—The rate of reaction of 2-hydroxypyridine with persulfate as a function of *p*H. Conditions: 0.025 *M* $S_2O_8^=$, 1.0 *M* 2-hydroxypyridine, (NaOH) + (NaCl) = 1.0 *M*, 29.9°.

4. Ionic Strength Dependence.—There is a positive and general salt effect whose magnitude is shown in Table III. It can be seen that sulfate ions, a product of the reaction, have no special effect. A plot of the log k *vs.* $\sqrt{\mu}$ is a straight line. However, the application of the Debye–Hückel equation to solutions of high ionic strength has no theoretical justification.[16]

TABLE II

PHENOL DEPENDENCE

Conditions: 1.7 *M* KOH and 40.0°

(2-OH pyridine) *M*	($S_2O_8^=$) *M*	k' (min.$^{-1}$)	k (l./mole-min.)
0.2	0.02	0.0425	0.2125
.1	.01	.0212	.2120
.05	.005	.01035	.2070
.02	.002	.00428	.2140

TABLE III

IONIC STRENGTH DEPENDENCE

	k (l./mole-min.)	μ
No additions	0.204^a	2
1 *M* NaCl	.268^a	3
$^1/_3$ *M* Na_2SO_4	.259^a	3
1 *M* NaBr	.269^b	3
1 *M* KCl	.277^b	3
1 *M* Na_2SO_4	.371^b	6

^{a} Conditions: 2 *N* NaOH, 0.2 *M* 2-OH pyridine, 0.024 *M* $S_2O_8^=$ and 40.0°. ^{b} Conditions: 2 *N* NaOH, 0.241 *M* 2-OH pyridine, 0.0175 *M* $S_2O_8^=$ and 40.0°.

5. pH Dependence.—At pH less than 8, the addition of 2-hydroxypyridine to a buffered solution of persulfate does not accelerate the spontaneous rate of decomposition of persulfate. At higher pH, the rate of disappearance of the persulfate increases and approaches a maximum value (Fig. 2, Tables IV and V). The shape of the rate *vs.* pH curve resembles the titration curve for 2-hydroxypyridine (p*K* 11.6).[17] This suggests that it is the phenolate ion and not the undissociated phenol which is the reactive species. Data on the pH dependence for the reaction with phenols of widely differing pK confirm this conclusion. Thus, salicylic acid (p*K* 13.4)[18] does not accelerate the spontaneous rate of disappearance of persulfate below pH 10. However, the reaction with *o*-nitrophenol (p*K* 7.2)[19] already proceeds at maximum velocity at pH 10, Table V.

TABLE IV

pH DEPENDENCE—2-HYDROXYPYRIDINE

Conditions: 0.2 *M* 2-OH pyridine, 0.02 *M* $S_2O_8^=$ and 40.0°

(NaOH) *M*	(NaCl) *M*	k' (min.$^{-1}$)
0.25	1.75	0.0341
0.50	1.50	.0371
1.0	1.0	.0371
2.0	..	.0408

TABLE V

pH DEPENDENCE—*ortho*-NITROPHENOL

Conditions: 0.197 *M* *o*-nitrophenol, 0.02 *M* $S_2O_8^=$ and 29.8°C.

pH	k' (min.$^{-1}$)
9.2	0.013^a
10	.016^a
10.9	.015^a
~14	.015^b

^{a} 0.06 *M* carbonate buffer. ^{b} 1.7 *M* KOH.

6. Lack of Inhibition by Oxygen and by Allyl Acetate.—There was no detectable effect on the rate of persulfate decomposition in reactions run with continuous nitrogen or oxygen bubbling.

The effect of allyl acetate on the Elbs reaction was most conveniently determined from the reaction of *o*-nitrophenol with persulfate. The high acidity of this phenol (p*K* 7.2) permitted the reaction to be run in the presence of allyl acetate at a pH (9.2) and a temperature (30°) at which the hydrolysis of allyl acetate was negligible.[13]

The free-radical nature of the persulfate oxidation of alcohols has been demonstrated by several investigators.[20] These reactions are inhibited by allyl acetate[13,21] which acts as a sulfate-radical trap. More recently, Ben-Zvi and Allen[22] have shown a similar inhibition by allyl acetate of the copper-catalyzed sulfate-radical oxidation of oxalate.

It was first determined that the oxidation of 2-propanol was inhibited under these conditions, which were somewhat different, particularly with regard to temperature from those used in previous

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[18] I. M. Kolthoff, *Rec. trav. chim.*, **42**, 969 (1923).

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[20] D. L. Ball, M. M. Crutchfield and J. O. Edwards, *J. Org. Chem.*, **25**, 1599 (1960). This paper has an excellent discussion of the persulfate oxidation of alcohols.

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[22] E. Ben-Zvi and T. L. Allen, *ibid.*, **83**, 4352 (1961).

studies.[13,21] A half-time of 157 min. was calculated for the initial rate of disappearance of 0.038 *M* persulfate in the presence of 1.015 *M* 2-propanol. In a parallel experiment, the addition of 0.1 *M* allyl acetate (final concentration) completely inhibited the disappearance of persulfate. Both reactions were carried out at 29.8° in 0.5 *M* potassium carbonate buffer, pH 9.2, under nitrogen.[20]

A series of experiments was carried out to determine the effect of allyl acetate on the persulfate oxidation of *o*-nitrophenol (Table VI). The concentrations of phenol and persulfate were varied over a tenfold range while the allyl acetate concentration was kept constant. The presence of allyl acetate had no detectable effect on the rate of the reaction.

TABLE VI

THE LACK OF EFFECT OF ALLYL ACETATE ON THE PERSULFATE OXIDATION OF *o*-NITROPHENOL

Conditions: 0.1 *M* allyl acetate, 0.6 *M* potassium carbonate buffer (*p*H 9.2) and 29.8°.

(Potassium *o*-nitrophenolate)	($S_2O_8^=$)	$t_{1/2}$ (min.)	*k* (l./mole-min.)
0.2	0.0205	52[a]	0.066[a]
.1	.0102	111[a]	.062[a]
.05	.00505	222[a]	.062[a]
.02	.00200	575[a]	.060[a]

[a] Identical values were obtained in the absence of allyl acetate.

7. Temperature Dependence.—The temperature dependence of the rate constant in the reaction with 2-hydroxypyridine is shown in Table VII.

TABLE VII

TEMPERATURE DEPENDENCE

Conditions: 2 *M* NaOH, 0.017 *M* $K_2S_2O_8$ and 0.0241 *M* 2-OH pyridine

Temp. (°C.)	*k* (l./mole-min.)
40.0	0.204
35.0	.117
29.9	.0810
24.8	.0506
18.5	.0280

An activation energy of 16.3 kcal./mole was calculated from a plot of log *k vs.* $1/T$.

Discussion

A mechanism which conforms to the experimental observations is a bimolecular ionic reaction in which the phenolate ion attacks the peroxy-bond of persulfate ion with displacement of a sulfate ion.

Cleavage of the persulfate molecule prior to attack of the phenolate anion is not consistent with the evidence presented in this study. Heterolysis, the decomposition of persulfate to sulfur tetroxide and sulfate, has been shown by Kolthoff and Miller[11] to be a significant reaction only in strongly acid solutions. These investigators demonstrated that heterolysis of persulfate was undetectable in 0.1 *M* NaOH; homolysis to sulfate radicals was the exclusive route of persulfate decomposition at high pH.

Sulfate radicals must certainly be produced under the conditions of the Elbs oxidation. Coupling products have been found[23] and in fact predominate under metal ion catalysis.[12] However, under the conditions we have described, the dissociation of persulfate to radical fragments must be quantitatively unimportant. The complete absence of allyl acetate inhibition is strong evidence against significant participation of sulfate radicals in the oxidation. In addition, the lack of any effect on the rate of the reaction by added cupric or ferrous ions or by oxygen makes a free-radical mechanism less likely. The kinetic evidence alone precludes homolysis or heterolysis as a unimolecular initiation step since neither mechanism predicts first-order dependence on phenolate concentration.

Preliminary experiments on the effect of structure of the phenol on reactivity have shown that the rate of consumption of persulfate is increased by electron-releasing substituents. Representative second-order rate constants are (*k*, l./mole-min. 30°, 2 *M* NaOH) 0.32 (salicylaldehyde), 0.96 (*o*-chlorophenol), 1.15 (phenol), 4.6 (*o*-cresol), 20.0 (guaiacol). However, since the stoichiometry has not yet been established for these reactions, conclusions drawn from them must be tentative. Furthermore, Bunnett, in his discussion of the isokinetic relationship,[24] has emphasized the danger of interpretation of rate data from a reaction series carried out at one temperature.

Acknowledgment.—We are indebted to Dr. B. M. Pitt for stimulating conversations.

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